Ch 8 Bonding / Ch 9 Hybridization

Everything around us is Bonded

Chemistry is Breaking -> Reforming Bonds

Graphite is smooth → slippery material
Diamond is ~~copper~~ hard and gemstone/cutting material → Different Bonds

8.1 Types of Bonds

↳ Form so system achieve lowest possible Energy

Ionic Bonds - form when atom that loses e^- Relatively easily reacts w/ atom that gains e^- (high e^- A) easily.

Ionic Compound -> formed from metal + nonmetal

Measure E of attraction b/w pairs of ions

Coulomb's Law $E = (2.31 \times 10^{-19}\ J \cdot nm)\left(\frac{Q_1 Q_2}{}\right)$

- E: Joules
- $Q_1 Q_2$: charge of ions
- distance B/w ionic centers

For NaCl

$$E = (2.31\times10^{-19}\ J\cdot nm)\left(\frac{(+1)(-1)}{0.276\ nm}\right) = -8.37\times10^{-19}\ J$$

Be aware may be in different units

(-) B/c get more stable pair has lower E

Calc Repulsion of 2 ~~cations~~ like ions

↳ will Be (+) B/c Repulsion

Easy to understand when

What about like elements (covalent bond)

↳ must overcome attraction + Repulsion of p+ e⁻

↳ Bond formed system goes to lowest possible E

Bond length - d w/ minimal PE / Bond E - E Req'd to Break Bond

Fig 8.1

Covalent Bond e⁻ shared by nuclei

In Between Ionic Bonds / Covalent Bonds

↳ Polar Covalent Bonds

Polar Covalent Bond → ~~non~~ metal

↳ not metal electron sharing.

uneven distribution of e⁻ $\overset{\delta+}{H}—\overset{\delta-}{F}$

↑ Polar Bond

Why? Electronegativity

8.2 ↳ ability of an atom to attract e⁻ in a molecule.

Linus Pauling First measured Relative E-neg

Show fig 8.3 - Periodic Trends pg 334 explain how e-neg calculated

Really difference in e-neg manages Bond character (compared: expected BE actual BE)

any difference gives some polarity do sample 8.1

Rule of Thm 2 nonmetals covalent
2 metal/nonmetal ionic

8.3

↳ Bond Polarity

dipole moment → separation of (+) + neg charge

$\overset{+}{\longrightarrow}$

HF

View w/ e⁻ static potential diagram

Red indicates e⁻ Rich areas

any polar 2 atom molecule will show dipole

polyatomic atoms too. H_2O + NH_3

polar Bonds / polar molecules

---

some molecules w/ polar Bonds But have polarities cancel out are non polar molecules

CO_2 → No way molecule lines up in electric field.

same for CCl_4 + SO_3 Samp 8.2

get fig do out.

8.5 Energy in Binary Comps (Ionic)

What factors influence ionic comp stability.

lattice E - The change in E that takes place when separated gaseous ions are packed together to form ionic Solid.

$$M^+_g + X^-_g \rightarrow MX_{(s)}$$

Show E changes w/ formation of $Li_{(s)} + \frac{1}{2}F_{2(g)} \longrightarrow LiF_{(s)}$

$Li_{(s)} \longrightarrow Li_{(g)}$ sublim	161 kJ
	520 kJ
$Li_{(g)} \longrightarrow Li^+_{(g)} + e^-$	
$\frac{1}{2}F_{2(g)} \longrightarrow F_{(g)}$	77
$e^- + F_{(g)} \longrightarrow F^-$	−328
$Li^+_{(g)} + F^-_{(g)} \longrightarrow LiF_{(s)}$	−1047 ← highly Exothermic Very Stable
	−617 kJ/mol

Ions arranged in crystal lattice → highly Repeating Pattern.

Show Fig 8.9, 8.10, 8.11

↑ Talk about in 2+ cations much higher lattice E. so more stable and more Favorably forms 2+ & 2− cations/anions

8.7 Covalent Bonds

↳ Bonds form to get lowest PE

when Break CH_4 into $4H + C$ Req 1652 kJ

or in Reverse E is Released

So Methane is more stable compared to Separate atoms

Typically we look @ Bond Energy of molecule w/ Respect to single Bonds

$$\frac{1652\ kJ}{4} = 413\ \frac{kJ}{mol}\ \text{C-H Bond}$$

↑ can use Bond energies to calc other Bond E's

For CH_3Cl → takes $1578\ \frac{kJ}{mol}$ to Break it down.

So $1\ C\text{-}Cl + 3\ (C\text{-}H)\ bonds = 1578\ kJ$

Plug in & solve for C-Cl = 339 kJ

8.8 Bond E + Chem Rxn

8.4 Bond E chart

Can use Bond E to calculate ΔH Rxn

$$H_2 + F_2 \longrightarrow 2HF$$

$$\Delta H = \sum D(\text{bonds Broken}) - \sum D(\text{Bonds formed})$$

1 Bond E; Σ year. Req.; Energy released

Calc for this ΔH_f HF = $-271\ \frac{kJ}{mol}$

TABLE 8.4 Average Bond Energies (kJ/mol)

Single Bonds						Multiple Bonds	
H—H	432	N—H	391	I—I	149	C=C	614
H—F	565	N—N	160	I—Cl	208	C≡C	839
H—Cl	427	N—F	272	I—Br	175	O=O	495
H—Br	363	N—Cl	200			C=O*	745
H—I	295	N—Br	243	S—H	347	C≡O	1072
		N—O	201	S—F	327	N=O	607
C—H	413	O—H	467	S—Cl	253	N=N	418
C—C	347	O—O	146	S—Br	218	N≡N	941
C—N	305	O—F	190	S—S	266	C≡N	891
C—O	358	O—Cl	203			C=N	615
C—F	485	O—I	234	Si—Si	340		
C—Cl	339			Si—H	393		
C—Br	276	F—F	154	Si—C	360		
C—I	240	F—Cl	253	Si—O	452		
C—S	259	F—Br	237				
		Cl—Cl	239				
		Cl—Br	218				
		Br—Br	193				

*C=O(CO_2) = 799

8.9 Bonding Model

Localized electron (LE) Model - molecule is composed of atoms that are bound together by sharing pairs of e⁻ using the atomic orbitals of the bound atoms

lone pair - e⁻ localized to atom

bonding pair - e⁻ shared B/w

→ 3 parts - Describe valence e⁻ w/ Respect to Lewis structures

Predicts molecular geometry VSEPR

Describes the atomic orbital to share e⁻ or lone pair

Draw Lewis structures to satisfy octet Rule

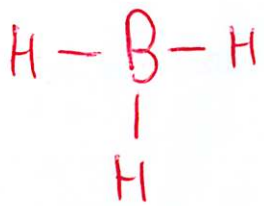
Remember draw ~~structures as~~ e⁻ pair —

Draw lewis structures for H_2O, CO_2, CN^-, NH_3, NO^+

Octet Rule Exceptions

Less than octet

Boron B/c 3 val e⁻



Reacts w/ free e⁻ pairs making a coordinate covalent bond

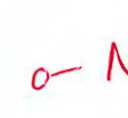
Exceeding octet Rule

SF_6 ← period 3

extra e⁻ are in d-orbitals

Write x-tures for PCl_5 I_3^- XeO_3 $BeCl_2$ ICl_4^-

Resonance more than 1 Lewis x-ture

NO_3^- 

Not 2 single > 3 Bonds w/ in B/w Strength + Length
1 double

electrons are delocalized b.c they can move around the x-ture

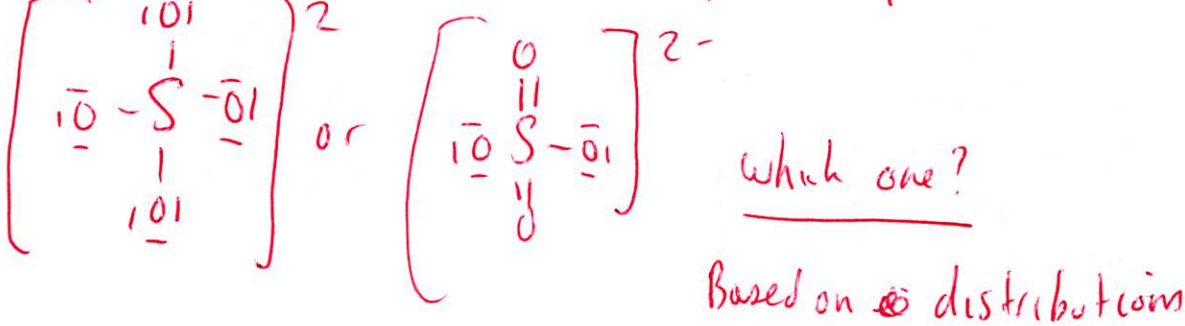
Even though we say e⁻ are localized they are Really Moving around. Res proves this → draw for NO_2^-

Free Radicals → odd e^- # NO

very Reactive.

Formal Charge – difference B/w # of e^- in free atom & # of val e^- assigned in atom in molecule

→ Why important b/c in polyatomic ions molecules there are many possible xtures. for example SO_4^{2-}



or

which one?

Based on ⊗ distributions of charge

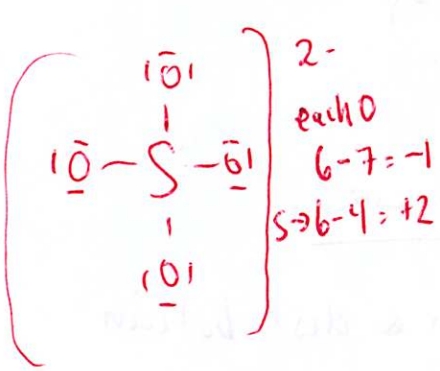
Can't use OX #'s b/c not Realistic assumptions of ⊗ charge distribution. They give all the e- to the more e-neg atom. So charges are exagerated.

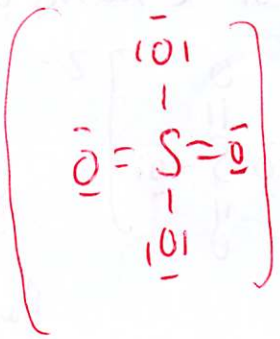
Formal chage helps out

FC = # of vale- on atom − # of vale- on atom in molec

↖ to get lone pair on atom
Belong to atom
shared e⁻ split evenly.

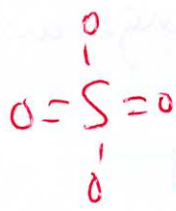
for Sulfate





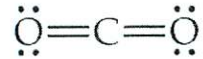
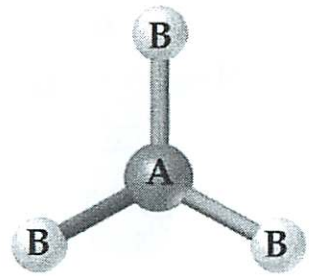
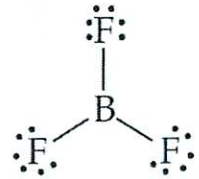
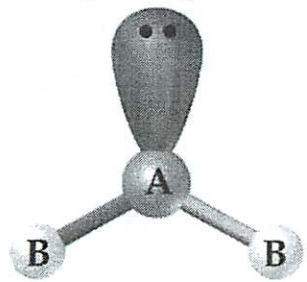
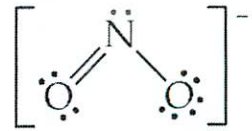
Structure assume atoms in molecules try to acheive charges as close to 0 as possible

(−) are going to reside on most e-neg atom

So  is Better + Resonance formate of

even more stable

TABLE 9.2 Electron-Domain Geometries and Molecular Shapes for Molecules with Two, Three, and Four Electron Domains Around the Central Atom

Number of Electron Domains	Electron-Domain Geometry	Bonding Domains	Nonbonding Domains	Molecular Geometry	Example
2	 Linear	2	0	 Linear	
3	 Trigonal planar	3	0	 Trigonal planar	
		2	1	 Bent	

From Brown / LeMay 10th Edition

TABLE 2 Electron-Domain Geometries and Molecular Shapes for Molecules with Two, Three, and Four Electron Domains Around the Central Atom

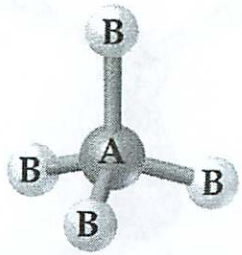
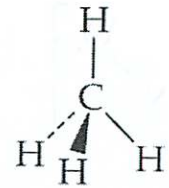
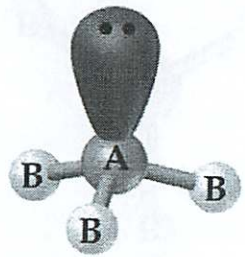
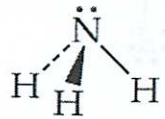
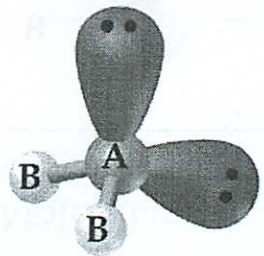
Number of Electron Domains	Electron-Domain Geometry	Bonding Domains	Nonbonding Domains	Molecular Geometry	Example
4	 Tetrahedral	4	0	 Tetrahedral	
		3	1	 Trigonal pyramidal	
		2	2	 Bent	

TABLE 9.3 Electron-Domain Geometries and Molecular Shapes for Molecules with Five and Six Electron Domains Around the Central Atom

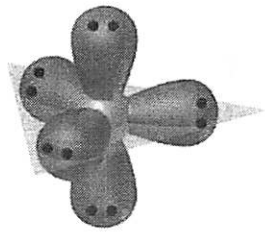
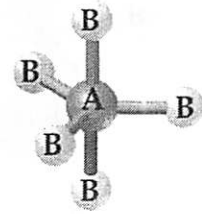
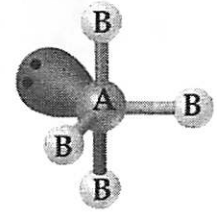
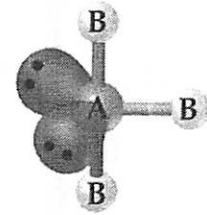
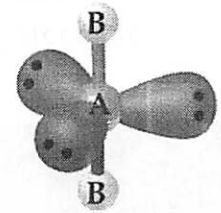
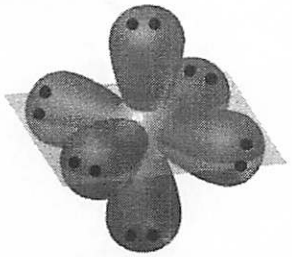
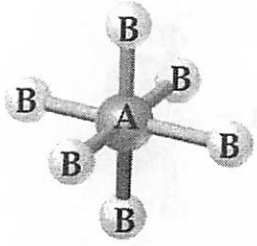
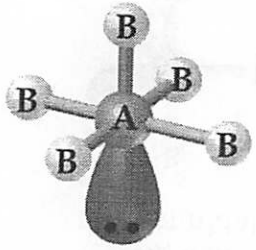
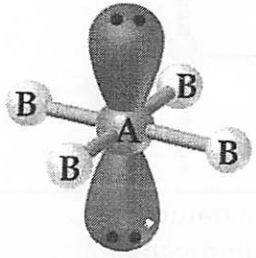
Total Electron Domains	Electron-Domain Geometry	Bonding Domains	Nonbonding Domains	Molecular Geometry	Example
5	 Trigonal bipyramidal	5	0	 Trigonal bipyramidal	PCl_5
		4	1	 Seesaw	SF_4
		3	2	 T-shaped	ClF_3
		2	3	 Linear	XeF_2

TABLE 9.3 Electron-Domain Geometries and Molecular Shapes for Molecules with Five and Six Electron Domains Around the Central Atom

Total Electron Domains	Electron-Domain Geometry	Bonding Domains	Nonbonding Domains	Molecular Geometry	Example
6	 Octahedral	6	0	 Octahedral	SF_6
		5	1	 Square pyramidal	BrF_5
		4	2	 Square planar	XeF_4

8:13 VSEPR - 3D shapes of molecules

↳ way to predict molecular x-tune

Figured by minimizing e⁻ pair Repulsion

Go over Bond angle each geometry.

why e⁻ Repulsion produces each Bond angle.

Lone pairs of e⁻ occupy more Room so the compress Bond ∠

Simply go over table and predict

Geometries from shapes

VSEPR shapes SO_2 / KrF_2 SO_3 NF_3

SF_4 XeF_4 IF_5

Ch 9

Hybridization - Explains what orbitals are doing in VSEPR

↳ B/c orbitals are of different energy.

So the orbitals fuse to make equal E orbitals

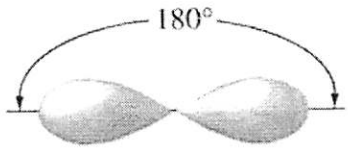
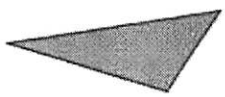
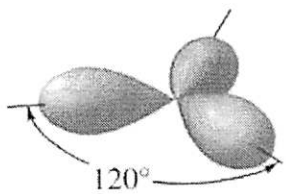
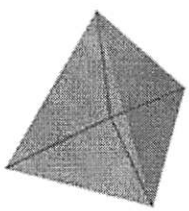
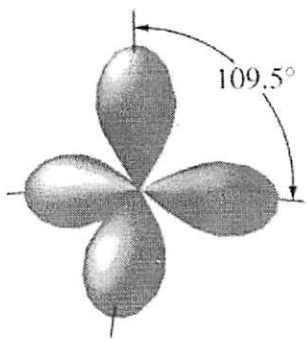
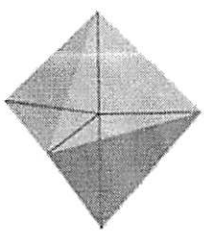
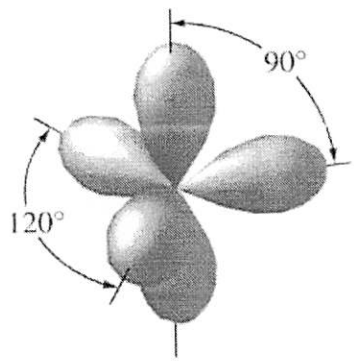
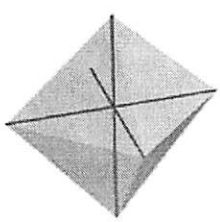
So s + p orbitals fuse to make hybrid orbitals

go over how they form / look

Bond L's, σ bond / π bond, sp^3 sp^2 sp

dsp^3 d^2sp^3

Look @ Geometries 9.24

Number of Effective Pairs	Arrangement of Pairs		Hybridization Required	
2		Linear	sp	
3		Trigonal planar	sp^2	
4		Tetrahedral	sp^3	
5		Trigonal bipyramidal	dsp^3	
6		Octahedral	d^2sp^3